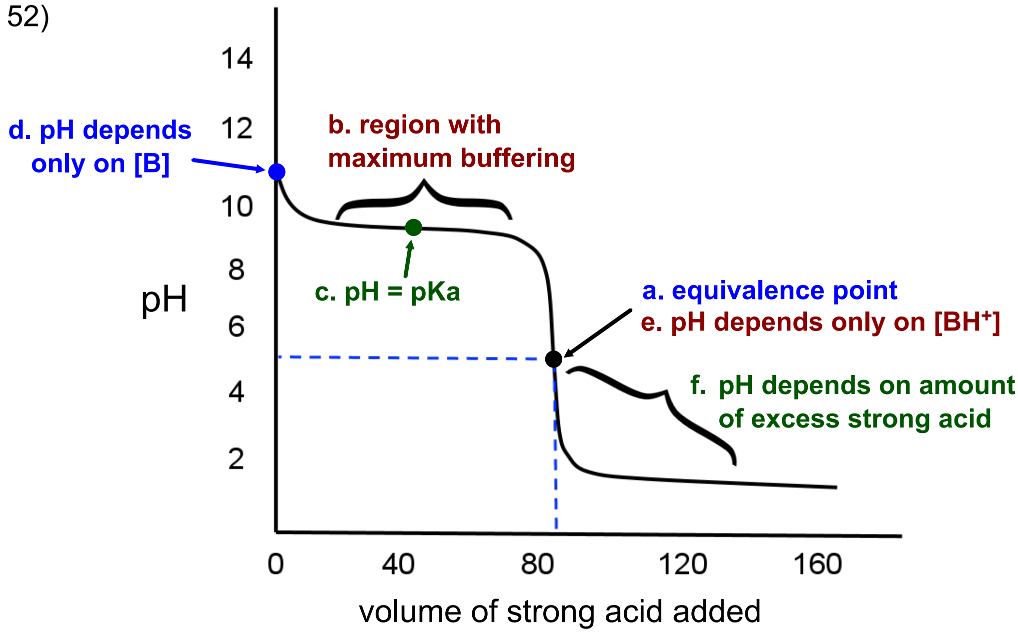
Assignment 15.2 Questions 52, 56, 57, 62, 66, 70

52)



56) a.

$$K_b = 3.0 \times 10^{-6} = \frac{[H_2NNH_3^+][OH^-]}{[H_2NNH_2]} = \frac{x^2}{0.100 - x}$$

$3.0 \times 10^{-6} (0.100) = x^2$ $\quad$ $pOH = -\log [5.48 \times 10^{-4}]$

$[OH^-] = x = 5.48 \times 10^{-4}$ M $\quad$ pOH = 3.26 $\quad$ **pH = 10.74**

b.

	$[H_2NNH_2]$ (mol)	$[H^+]$ (mol)	$[H_2NNH_3^+]$ (mol)
I	0.0100	0.0040	0
C	- 0.0040	- 0.0040	+ 0.0040
E	0.0060	0	0.0040

$$pOH = pKb + \log \frac{(acid)}{(base)} = 5.52 + \log \frac{(0.004)}{(0.006)} = 5.35$$ **pH = 8.65**

c.

	$[H_2NNH_2]$ (mol)	$[H^+]$ (mol)	$[H_2NNH_3^+]$ (mol)
I	0.0100	0.0050	0
C	- 0.0050	- 0.0050	+ 0.0050
E	0.0050	0	0.0050

$$pOH = 5.52 + \log \frac{(0.005)}{(0.005)} = 5.52$$ **pH = 8.48**

56) d.

	$[H_2NNH_2]$ (mol)	$[H^+]$ (mol)	$[H_2NNH_3^+]$ (mol)
I	0.0100	0.0080	0
C	- 0.0080	- 0.0080	+ 0.0080
E	0.0020	0	0.0080

$$pOH = 5.52 + \log \frac{(0.008)}{(0.002)}$$

$pOH = 6.12$

pH = 7.88

e.

	$[H_2NNH_2]$ (mol)	$[H^+]$ (mol)	$[H_2NNH_3^+]$ (mol)
I	0.0100	0.0100	0
C	- 0.0100	- 0.0100	+ 0.0100
E	0	0	0.0100

$$[H_2NNH_3^+] = \frac{0.0100 \text{ mol}}{0.150 \text{ L}}$$

$[H_2NNH_3^+] = 0.0667$ M

$$K_b = 3.3 \times 10^{-9} = \frac{[H_2NNH_2]\,[H^+]}{[H_2NNH_3^+]} = \frac{x^2}{0.0667}$$

$3.3 \times 10^{-9}\ (0.0667) = x^2$

$[H^+] = x = 1.49 \times 10^{-5}$ M

pH = 4.83

f.

	$[H_2NNH_2]$ (mol)	$[H^+]$ (mol)	$[H_2NNH_3^+]$ (mol)
I	0.0100	0.0200	0
C	- 0.0100	- 0.0100	+ 0.0100
E	0	0.0100	0.0100

$$[H^+] = \frac{0.0100 \text{ mol}}{0.200 \text{ L}}$$

$[H^+] = 0.0500$ M

pH = 1.301

With the concentration of H^+ being so high, at this point the concentration of the weak acid, $H_2NNH_3^+$, doesn't make a significant difference.

57) a. $K_a = 10^{-3.86} = 1.38 \times 10^{-4}$

$$1.38 \times 10^{-4} = \frac{[L^-]\,[H^+]}{[HL]} = \frac{x^2}{0.100 - x}$$

$1.38 \times 10^{-4}\ (0.100) = x^2$

$[H^+] = x = 0.00365$ M

pH = 2.43

or 2.44 using quadratic equation

b.

	[HL] (mol)	$[OH^-]$ (mol)	$[L^-]$ (mol)
I	0.0025	0.0004	0
C	- 0.0004	- 0.0004	+ 0.0004
E	0.0021	0	0.0004

$$pH = pKa + \log \frac{(\text{base})}{(\text{acid})} = 3.86 + \log \frac{(4 \text{ parts})}{(21 \text{ parts})} = \mathbf{3.14}$$

Until we reach the equivalence point, we can keep using this equation, just with different ratios of weak acid to conjugate base.

$$pH = 3.86 + \log \frac{(8 \text{ parts})}{(17 \text{ parts})} = \mathbf{3.53}$$

$$pH = 3.86 + \log \frac{(12.5 \text{ parts})}{(12.5 \text{ parts})} = \mathbf{3.86}$$

$$pH = 3.86 + \log \frac{(20 \text{ parts})}{(5 \text{ parts})} = \mathbf{4.46}$$

$$pH = 3.86 + \log \frac{(24 \text{ parts})}{(1 \text{ parts})} = \mathbf{5.24}$$

$$pH = 3.86 + \log \frac{(24.5 \text{ parts})}{(0.5 \text{ parts})} = \mathbf{5.55}$$

$$pH = 3.86 + \log \frac{(24.9 \text{ parts})}{(0.1 \text{ parts})} = \mathbf{6.26}$$

At the equivalence point:

	[HL] (mol)	[OH^-] (mol)	[L^-] (mol)
I	0.0025	0.0025	0
C	- 0.0025	- 0.0025	+ 0.0025
E	0	0	0.0025

$$[L^-] = \frac{0.0025\text{mol}}{0.050\text{ L}}$$

$[L^-] = 0.050$ M

$$K_b = 7.25 \times 10^{-11} = \frac{[HL]\,[OH^-]}{[L^-]} = \frac{x^2}{0.050}$$

$x = [OH^-] = 1.90 \times 10^{-6}$ M

pOH = 5.72

pH = 8.28

After the equivalence point, the pH is determined by the excess of OH- after the lactic acid has been neutralized.

$[OH^-] = \mathbf{0.00251} - 0.00250 = \frac{0.00001\text{mol}}{0.0501\text{ L}} = 0.0002$ M

pOH = 3.7
pH = 10.3

$[OH^-] = \mathbf{0.0026} - 0.0025 = \frac{0.00010\text{mol}}{0.051\text{ L}} = 0.0020$ M

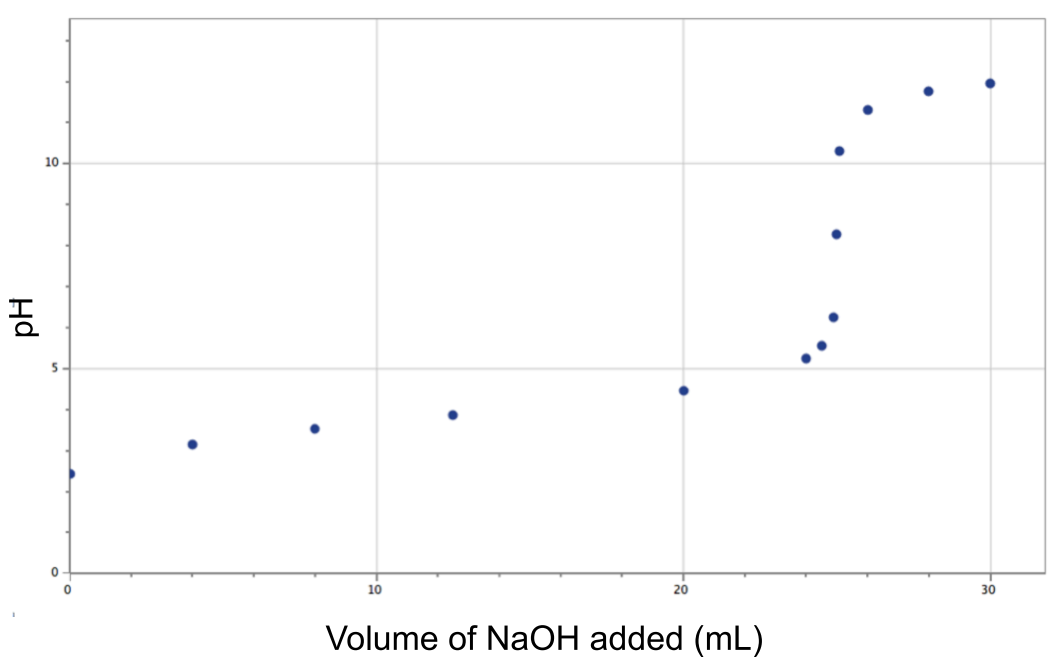
pOH = 2.70
pH = 11.30

$[OH^-] = \mathbf{0.0028} - 0.0025 = \frac{0.00030\text{mol}}{0.053\text{ L}} = 0.0057$ M

pOH = 2.25
pH = 11.75

$[OH^-] = \mathbf{0.0030} - 0.0025 = \frac{0.00050\text{mol}}{0.055\text{ L}} = 0.0091$ M

pOH = 2.04
pH = 11.96



62)

	$[CH_3NH_2]$ (mol)	$[H^+]$ (mol)	$[CH_3NH_3^+]$ (mol)
I	0.050	0.025	0
C	- 0.025	- 0.025	+ 0.025
E	0.025	0	0.025

$$pOH = pKb + \log\frac{(acid)}{(base)} = 3.36 + \log\frac{(0.025)}{(0.025)} = 3.36 \qquad \mathbf{pH = 10.64}$$

	$[CH_3NH_2]$ (mol)	$[H^+]$ (mol)	$[CH_3NH_3^+]$ (mol)
I	0.050	0.050	0
C	- 0.050	- 0.050	+ 0.050
E	0	0	0.050

$$[CH_3NH_3^+] = \frac{0.050 \text{ mol}}{0.150 \text{ L}}$$

$[CH_3NH_3^+] = 0.333$ M

$$K_a = 2.3 \times 10^{-11} = \frac{[CH_3NH_2]\,[H^+]}{[CH_3NH_3^+]} = \frac{x^2}{0.333}$$

$2.3 \times 10^{-11}\,(0.333) = x^2$

$[H^+] = x = 2.75 \times 10^{-6}$ M

pH = 5.56

66) An indicator is useful if it changes color when pH = pKa (+1 or -1)

pH = pKa = -log (5.0×10^{-6}) = 5.30 (changes between 4.3 - 6.3)

a. If methyl red is used for the titration of a weak acid with a strong base, the indicator will begin to change from red to yellow at a pH of 4.3, but the acid will not be neutralized until a pH above 7, so it would not be an effective indicator.

b. If used for the titration of a weak base with a strong acid, the indicator will begin to change color at a pH of 6.3. This may be close to the equivalence point, depending on the pK_b of the weak base, so it might be an effective indicator.

70) The equivalence point for Problem #56 was 4.83, so we need an indicator that changes color in that range. Bromcresol Green would be the most effective indicator.